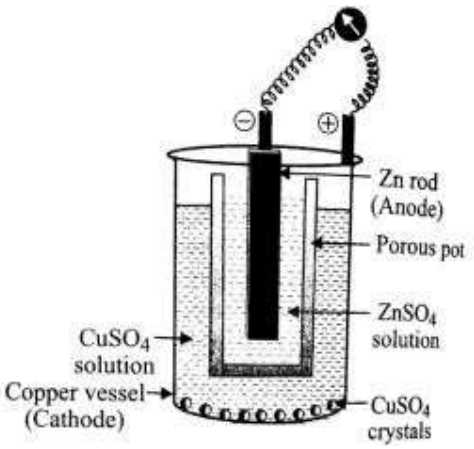
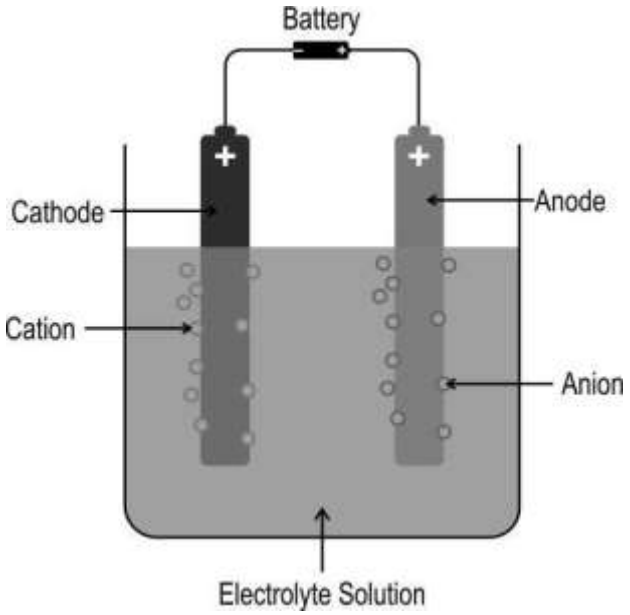
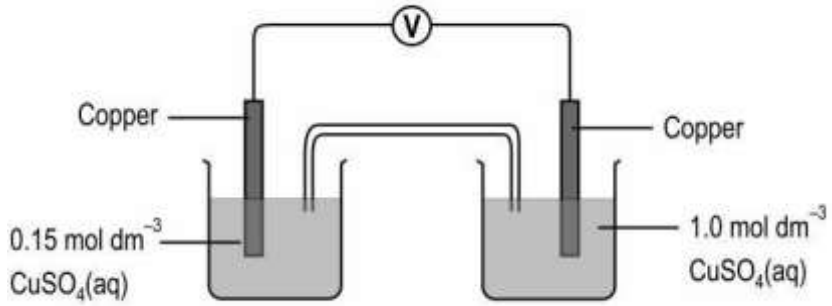
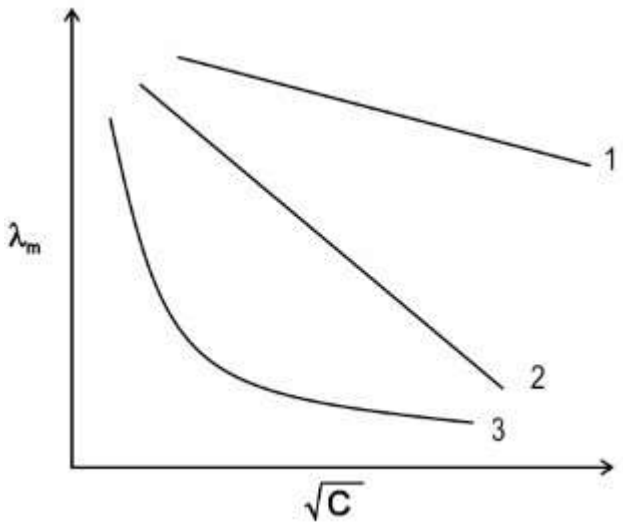
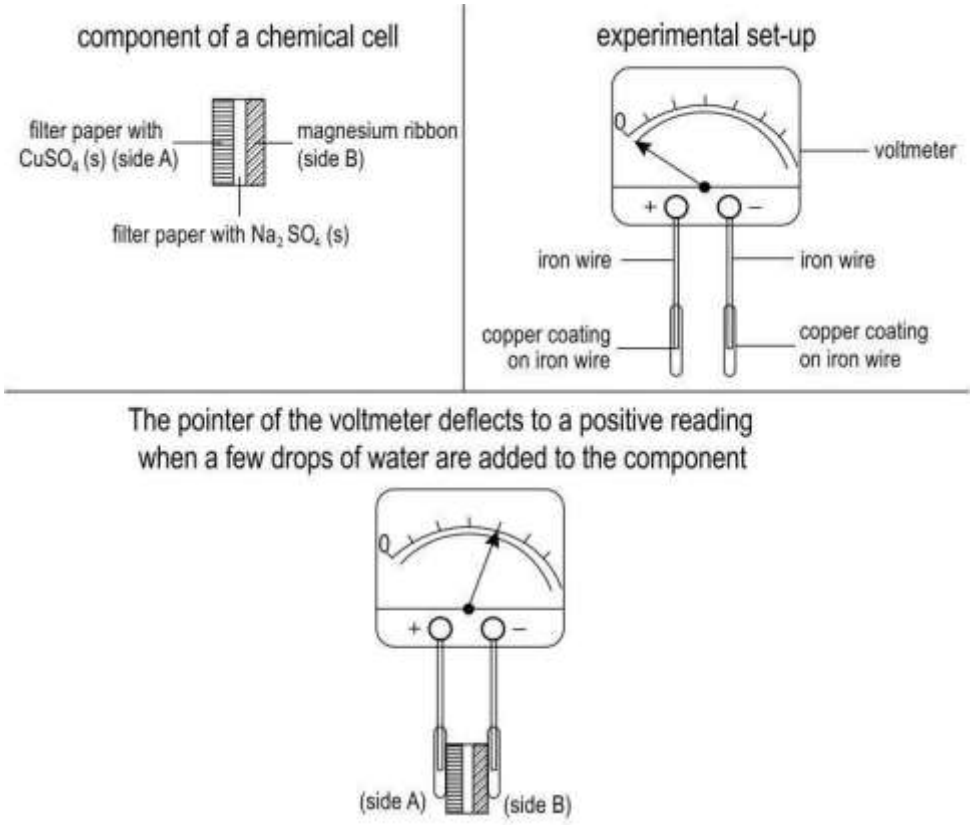
3. ELECTROCHEMISTRY

Q.No	Question	Marks
	Multiple Choice Question	
Q.49	If the standard emf of Galvanic cell–I: $Cu_{(s)}/Cu^{2+}_{(aq)}$II $Ag^{+}_{(aq)}/Ag_{(s)}$ is 0.46V, and the standard emf of Galvanic cell–II: $3\ Cu_{(s)} + 6\ Ag^{+}_{(aq)} \rightarrow 3Cu^{2+}_{(aq)} + 6\ Ag\ (s)$ is 0.46q V. What is the value of q? A. 3 B. 2 C. 1 D. Infinite	1
Q.50	The electrochemical cell made up of Zn and Cu half-cell is called Daniell cell. The emf of a Daniell cell is 1.10V.  When an external voltage greater than 1.10 V is applied to this cell, which of the following change will be observed in the cell? A. Zn electrode will act as an anode. B. Current will flow from Cu half cell to Zn half cell. C. Electrochemical cell continue to work fast. D. Cell will act as electrolytic cell.	1
Q.51	An electrolytic cell has an anode and cathode made up of graphite. At the anode, Cl_2 gas is released and at the cathode, H_2 gas is released. Which of the following electrolytes in the cell can produce these gases? A. $NH_4Cl_{(aq)}$	1

	B. Molten NH_4Cl C. NaCl (aq) D. Molten NaCl	
Q.52	There are two beakers 'A' and 'B' containing KCl and CH_3COOH solutions respectively. On adding water to beakers A and B, which of the following change in Λ_m of the solutions will be correct? A. It increases sharply in beaker A and slowly in beaker B B. It increases slowly in beaker A and sharply in beaker B C. It decreases in beaker A but no change in beaker B. D. There is no change in beaker A but it decreases slowly in beaker B.	1
Q.53	Copper metal is purified by electrolytic refining. If the electrolyte used for refining of copper in an electrolytic cell is aq.salt solution of copper, which out of the following statement about this cell is INCORRECT? A. The impure Copper rod undergoes oxidation. B. Oxidation takes place at the anode. C. Impure copper rod acts as the negative electrode. D. Pure copper rod acts as a cathode.	1
Q.54	Given that the standard reduction potential for Fe^{3+}/Fe^{2+}= 0.77 V and I_2/I^-= 0.54V. Which of the following is correct when the cell is made by using Fe^{3+} and I^- salt solutions? A. The standard emf of the cell is −0.23 V B. The standard emf of the cell is +0.23 V C. The standard emf of the cell is 1.31 V D. The standard emf of the cell is -1.31 V	1
Q.55	Under which of the following conditions will the chemical reaction in an electrochemical cell will be spontaneous? A. E^0_{cell} = +ve, ΔG= +ve B. E^0_{cell} = −ve, Δ G= −ve C. E^0_{cell} = +ve, Δ G= −ve D. E^0_{cell} = −ve, ΔG= +ve	1

Q.56	How much electricity in Faraday is required for the complete reduction of MnO_4^- ions present in 500 ml of 0.5 M solution to Mn^{2+}? A. 5 F B. 2.5 F C. 2.25 F D. 1.25 F	1
Q.57	The image below shows electrolysis of an electrolyte using a DC voltage source.  Based on this, Which of the following statements is/are correct? (i) The solution remains electrically neutral during electrolysis. (ii) Electrons flow from the current source towards the solution at one electrode, and an equal number of electrons flow away from the solution at the other electrode. (iii) The number of positive ions moving towards one electrode is always equal to the number of negative ions moving towards the other electrode. A. i only B. i and ii only C. ii and iii only D. all- i, ii, and iii	1
Free Response Question/ Subjective Type		
Q.58	In the chemistry lab, Zoya set up an electrochemical cell as shown below:	2

	 At room temperature, she found that the initial voltmeter reading was +0.16v. (i) The standard electrode potential for the Cu^{2+}/Cu electrode is given by $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$; $E^o = +0.34$ V Calculate the electrode potential of the electrode on the left-hand side of the above electrochemical cell. (ii) Explain the importance of the salt bridge between the two solutions.	
Q.59	(i) The molar conductivity vs √c curve for NaCl, HCl, and NH_4OH are shown below in random order.  Identify which graph corresponds to HCl, NaCl, and NH_4OH. (ii) Give reasons to justify your answer in (i).	3
Q.60	The value of E^0 for two half cells are given as: $O_2 + 4e^- + H^+ \rightarrow 2H_2O$; $E^0 = 1.228$V and, $E^0Ce^{4+}/Ce^{3+} = 1.60$ V	4

	Based on the above data, will Ce ion undergo oxidation or reduction in water? Explain it with the help of overall cell reaction. Also, find the emf of the cell.	
Q.61	There are four electrodes A, B, C, and D. E^0 values of the electrodes are as follows. **Electrodes** / **Electrode Potential** A / $A/A^- = 0.96$ V B / $B^-/B^{2-} = -0.12$V C / $C^+/C = 0.18$V D / $D^{2+}/D = -1.12$V The combination of which of two electrodes will give the largest cell potential? Justify your answer. Also, find the emf of the cell.	3
Q.62	The diagrams below show the component of a chemical cell, an experimental set-up, and how the pointer of the voltmeter deflects when the set-up is connected to the component. Note that in the chemical cell, a filter paper with Na_2SO_4 acts as a salt bridge.  	2

	Why does the pointer of the voltmeter deflect as shown when a few drops of water are added to the component?	
Q.63	In a galvanic cell when the potential difference becomes zero, the cell is said to be in an equilibrium state. Establish the relation between E^0 and equilibrium constant at 298 k in a Daniell cell. The E^0 value of the Daniell cell is 1.10V. *(R= 8.314 j/k/mol, F = 96500 C)*	3
Q.64	i) Write down the complete cell reactions taking place at anode and cathode in a zinc/carbon dry cell. ii) Is the above given cell a primary cell or a secondary cell? Explain.	2
Q.65	How much time does it require to reduce 3 moles of iron (III) to 3 moles of iron (II) ion by passing a 2.0 amp current? *(Note: For calculations use 1 Faraday = 96500 Coulombs.)*	2
Q.66	A rusted piece of iron undergoes electrochemical reactions. Write the chemical reactions taking place at the following spots of that rusting piece of iron: a) At the spot that behaves as an anode b) At the spot that behaves as a cathode c) The overall balanced chemical reaction d) Further oxidation of ferrous ion into rust	2
Q.67	The Gibbs energy change for the reduction of Al_2O_3 at 500°c is given as: $2/3\ Al_2O_3$ ----- > $4/3\ Al^3 + O_2$; $\Delta G = +960KJ$ Calculate the minimum potential difference required to reduce 2/3 mole of Al_2O_3 at 500°C. (*1F = 96500C)*	3
Q.68	In an experiment, the electrolysis of copper sulfate solution takes place under the following conditions- - Electrolysis time (t) = 10 min. - Current passed (I) = 1.5 amp. What mass of copper will be deposited at the cathode in this experiment? *(Note: atomic mass Cu= 63.5g; For calculation use 1 Faraday = 96500 Coulombs.))*	3

Q.69	The electrolytic conductivity of $BaCl_2$ solution is 0.580 Sm^{-1}. Find out molar concentration of the solution if molar conductivity of this solution is 2.416×10^{-2} Sm^2/mol.	2
Q.70	The molar conductivity of a dilute solution of methanoic acid is 46.1 S cm^2/mol. Calculate its degree of dissociation. *(Given $\lambda^0(H^+)$ = 349.6 S cm^2/mol and $\lambda^0(HCOO^-)$ = 54.6 S cm^2/mol)*	2
Q.71	Two electrolytic cells A and B containing electrolytic solutions of $CuSO_4$ and $AgNO_3$ respectively are connected in a series. A steady current of 1.5 amperes is passed through them. Based on this information, answer the following questions. a) Find out the time 't 'in minutes required to deposit 1.34 g of the silver in cell 'B'. b) What mass of copper will be deposited at the cathode of cell A under the same experimental condition. *(Given - Atomic mass Cu=63.5 g, Ag= 108 g and 1F= 96500 C)*	3
Q.72	Given is an electrochemical cell; $Mg/Mg^{2+}_{(aq)}$ || $Cu^{2+}_{(aq)}/Cu_{(s)}$ Calculate the equilibrium constant of the cell at 25°C when the emf of the cell is zero. *(E^0 Mg^{2+}/Mg= -2.37V, Cu^{2+}/Cu= 0.34V, 2.303RT/F= 0.0591)* *Use log and antilog table if needed.*	2
Q.73	For an experiment, Aman prepared a 1-litre $FeSO_4$ solution of 1 M concentration and stored the solution in a glass jar. Before starting the experiment, Aman wants to stir the solution. Which of the following spoons should he use for this purpose and why? Aluminium spoon (Al^{3+}/Al = -1.66V) Copper spoon (Cu^{2+}/Cu = 0.34V) (*Given: E^0/V Fe^{2+}/Fe = -0.44V*)	2
Q.74	The potential of Zn, Cu and Ag half cells are given below; $Zn_{(s)}$ -----> $Zn^{2+}_{(aq)}$ + $2e^-$; E^0= +0.76V $Cu_{(s)}$ ------> $Cu^{2+}_{(aq)}$ + $2e^-$; E^0= -0.34V $Ag_{(s)}$ --------> $Ag^+_{(aq)}$ + $1e^-$; E^0 = -0.80V. Using the data above given, i) Write the correct cell representation of a cell with a cell potential equal to 0.46V.	2

	ii) Calculate the value of standard free energy change (ΔG^0) for the cell above given. (F = 96500 C/mol)	
Q.75	Predict the feasibility of the following reaction. Justify your answer. $Ag(s) + Fe^{3+}_{(aq)} \text{------>} Ag^{+}_{(aq)} + Fe_{(s)}$ (Given: $Ag^{+}_{(aq)}/Ag_{(s)} = 0.80V$, $Fe^{3+}/Fe_{(s)} = 0.77V$)	2
Q.76	In a Standard Hydrogen Electrode (SHE), the platinum wire is normally dipped in 1 M con. HCl solution. Find out the potential of SHE if the platinum wire is dipped in a solution containing 1×10^{-10} M H^+ concentration.	2
Q.77	Calculate the charge required in coulombs to reduce 0.5 moles of $Cr_2O_7^{2-}$ ion to Cr^{3+} in acid solution. *(1 Faraday = 96500C)*	2
Q.78	One Faraday of electric charge is passed through the electrolytic cells placed in a series containing solution of Ag^+, $Cu^{2+}$and Al^{3+}respectively.Find out the simple mass ratio of the metals deposited at the respective electrodes. (Given - Atomic mass Ag=108g, Cu=63.5g. Al=27g)	2
Q.79	Imagine you are in a chemistry lab and the teacher is explaining the electrolysis of $CuSO_4$ solution and the products liberated after electrolysis. The teacher made two Setups for the electrolysis process. In Set up-i electrolysis of $CuSO_4$ solution is done by using Pt electrodes and in Set up-II electrolysis of $CuSO_4$ solution is done by using Cu electrodes. Answer the following questions based on this: i) In which Set up I or II will the colour of $CuSO_4$ solution fades away and why? ii) Write the chemical reaction taking place at the Cu anode in Set up II. iii) Name the product obtained at the anode in Set up I. iv) Which out of Set up I or II depict refining of crude copper?	3

Answer Key & Marking Scheme

Q.No	Answers	Marks
Q.49	C. 1	1
Q.50	D. Cell will act as electrolytic cell.	1
Q.51	C. NaCl (aq)	1
Q.52	B. It increases slowly in beaker A and sharply in beaker B	1
Q.53	C. Impure copper rod acts as the negative electrode.	1
Q.54	B. The standard emf of the cell is +0.23 V	1
Q.55	C. E^0_{cell} = +ve, Δ G= −ve	1
Q.56	D. 1.25 F	1
Q.57	B. i and ii only	1
Q.58	(i) Electrode potential of the electrode on the left-hand side is given by: => $E^o_{cell} = E^o_{cathode} - E^o_{anode}$ => $0.16 = 0.34 - E^o_{anode}$ => E^o_{anode} = +0.18 V (ii) It allows mobile ions to move through it between the solutions and maintain the charge balance.	2
Q.59	(i) From the above graph, 1 corresponds to HCl 2 corresponds to NaCl 3 corresponds to NH_4OH (ii) Explanation: - When the above compounds dissociate, H+ has the highest mobility in comparison with Na, because the Molar mass of H+ is less than Na+ ion. - HCl and NaCl are strong electrolytes compared to NH_4OH which is a weak base. [1 mark]	3

	- Strong electrolytes are already completely dissociated and there is a small increase (change) in dissociation on dilution. For weak electrolytes, the degree of dissociation increases to a greater extent/abruptly and follows the non-linear curve. - so at a given concentration, molar conductivities of HCl>NaCl>NH_4OH [1 mark]	
Q.60	$Ce^{4+} + e^- \rightarrow Ce^{3+}$; $E^0 = 1.60V$ ...(i) $O_2 + 4e^- + 4H^+ \rightarrow 2H_2O$; $E^0 = 1.228V$ ...(ii) Since the reduction potential of the Ce half cell is more than that of water, Ce will undergo reduction. [1 mark] **Explanation using reaction and emf of cell:** Multiplying eq(i) by 4: $4Ce^{4+} + 4e^- \rightarrow 4Ce^{3+}$; $E^0 = 1.60V$ Reversing eq(ii): $2H_2O \rightarrow O_2 + 4e^- + 4H^+$; $E^0 = -1.228V$ Adding the two eqs give: $2H_2O + 4Ce^{4+} \rightarrow O_2 + 4H^+ + 4Ce^{3+}$; $E^0 = 0.372V$; [1 mark] Since E^0 is positive, the reaction is spontaneous. Ce^{4+} will undergo reduction. [1 mark]	4
Q.61	A combination of electrodes A and D will give the largest cell potential. **Justification**: - Since, $E_{cell} = E_{red} - E_{ox}$ - For the largest cell potential, we need an electrode with very high (+ve) reduction potential at the cathode and another electrode with very low (-ve) reduction potential at the anode. [1 mark] => Electrode A/A^- = 0.96V shows highest reduction potential; Electrode D^{2+}/D = -1.12V shows least reduction potential => emf of the cell = 0.96 - (-1.12) V = 2.08 V [1 mark]	3
Q.62	- When a few drops of water are added, it acts as an electrolyte solution to form an electrochemical cell. [0.5 marks] - $CuSO_4$ ionizes to form Cu^{2+} and SO_4^{2-} ions. [0.5 marks] - Mg ionises to $Mg^{2+} + 2e^-$ [0.5 marks]	2

	- The presence of free electrons in the cell gives rise to emf and hence the pointer deflects. [0.5 marks]	
Q.63	Nernst equation (Daniell cell) : $E_{cell} = E^0_{cell} - 2.303\ RT/nF \log [Zn(aq.)^{2+}]/[Cu_{(aq.)}^{2-}]$; where R=8.314 j/k/mol F=96500 Cmol^{-1} T= 298 K - On substituting value $E_{cell} = E_{cell}^0 - 0.059/n \log [Zn^{2+}]/[Cu^{2+}]$ [1 M] - At equilibrium E_{cell}=0,and $[Zn^{2+}/Cu^{2+}] = \log k_c$ Then, $E^0_{cell} = (0.059/n) \times \log k_c$ [1] - For Daniell cell => E_{cell}^0= 1.10V, n=2 So, $1.10 = (0.059/2) \times \log K_c$ => $\log k_c = 2.20/0.059$ [1] or =>$\log k_c$=37.28	3
Q.64	i)In a Zinc/Carbon dry cell complete cell reaction is: $Zn_{(s)} + 2MnO_{2(s)} + 2 NH_4^+(aq) \rightarrow Zn^{2+}(aq) + Mn_2O_{3(s)} + 2NH_{3(aq)} + H_2O_{(l)}$ or $Zn_{(s)} + MnO_{2(s)} + NH_4^+ \rightarrow Zn^{2+}(aq) + MnO(OH)_{(s)} + NH^3(aq)$ *(give 1 mark for any)* ii) The Zinc/Carbon dry cell is a primary cell. - A primary cell is one in which redox reaction cannot be reversed. The Zinc/Carbon cell becomes dead after a long time of use i.e.it stops working. This shows it is a primary cell. *(give 0 marks if explanation not given)*	2
Q.65	As Q = I x t ...(i) where Q= charge (coulomb) I = current(amp) t = time(sec) Required equation:	2

	$3Fe^{3+} + 3e$ --- > $3Fe^{2+}$ charge required is = 3 F Substituting the value in eq.(i) => 3 x 96500 C = 2 amp. x t t = 144750 s Or t = 2412.5 minutes Or t = 40.21 hrs	
Q.66	Chemical reactions are as follows: a) At anode: $Fe_{(s)}$ ---> $Fe^{2+}{}_{(aq)} + 2e^-$ b) At cathode: $O_2(g) + 4H^+{}_{(aq)} + 4e^-$ --> $2H_2O_{(l)}$ c) Over all reaction: $2Fe_{(s)} + O_{2(g)} + 4H^+{}_{(aq)}$ --> $2Fe^{2+}(aq) + 2H_2O_{(l)}$ d) Further oxidation : $2Fe^{2+}{}_{(aq)} + 2H_2O_{(l)} + 1/2O_{2(g)}$ --> $Fe_2O_3{}_{(s)} + 4H^+{}_{(aq)}$	2
Q.67	Given is : $\Delta G = +960 \times 10^3$ J ; F = 96500 C Required formula: $\Delta G = - n E F$ Calculating n for 2/3 moles of Al_2O_3 $Al_2O_3 > 2Al^{3+} + 3/2\ O_2$ $(2\ Al^{3+} + 6e > 2Al)$ $2/3\ Al_2O_3 > 4/3\ Al + O_2$ ∵ 1 mole Al_2O_3 is reduced by 6 mol e^- ∴ 2/3mole Al_2O_3 is reduced by e^- = 6x2/3= $4e^-$ Substituting the value in formula 960×10^3 J = - 4 x E x 96500 C E= (- 960×10^3 J/ 4 x 96500 C) E = - 2.487 V Hence minimum potential difference required= 2.487V OR ≈ 2.5 V	3
Q.68	Charge=current x time Q = I x t	3

	Q = 1.5 amp x 10 x 60s Q = 900 C Calculating charge and amount of copper For 1 mole Cu^{2+}ion Cu^{2+}+ 2e ----> Cu 2F 63.5 [0.5 marks] - Charge required to deposit 1 mol copper = (2 x 96500 C) [0.5 marks] ∵ 2 x 96500 C charge deposit mass of Cu = 63.5gr. ∴ 900 C charge will deposit mass of Cu = (63.5 g x 900 C)/(2 x 96500) C [0.5 marks] Amount of Cu at cathode = 0.296 g [0.5 marks]	
Q.69	Step-I Calculation of Concentration $\Lambda_m = k/C$ Where; k= electrolytic conductivity C=molar concentration Λ_m=molar conductivity Substituting values; $C = 0.580/2.416x10^{-2}$ $C = 0.24x10^2 mol/m^3$ Step II molar concentration $C= 0.24x10^2 mol/m^3$ $(1m^3=1000L)$ $C= 0.24x10^2 mol/1000L$ C= 0.0240 mol/L	2
Q.70	Calculating the degree of dissociation: $\alpha = \Lambda_m / \Lambda_m^0$ $\Lambda_{m = 46.1}$ Λ_m^0 = ? - Calculating Λ_m^0 using Kohlrausch law $\Lambda_m^0 = \lambda(H)^+ + \lambda(HCOO)^-$	2

	Λ_m^0 = 349.6 + 54.6 Λ_m^0 = 404.2 S cm ^{2}/mol $\alpha = \Lambda_m / \Lambda_m^0$ α = 46.1/404.2 =0.114 Degree of dissociation = 0.114	
Q.71	a) Step-I. Calculating charge available Ag^+ (aq) + 1e-----> $Ag_{(s)}$ (1mol e^- = 1F charge) (Atomic mass of 1 mole of Ag = 108g) ∵ 108 gr. of Ag is deposited by 1F electric charge ∴ 1.34 gr. of Ag is deposited by electric charge = 96500/108 x 1.34 =1260 C (0.5marks) Step-II. Calculating time 't' for deposition of Ag We know - Q = I X t (0.5) Thus t= Q/I = 1260C/1.5amp t = 840 seconds. (0.5) In minutes; 840/60 = 14 minutes. (0.5) b) Calculating mass of Cu deposited in cell A. Cu^{2+} (aq.) + 2e-------> Cu(s) (2 mole e^- = 2 F charge) (mass of 1 mole of Cu = 63.5gram) We know- ∵ 96500 X 2 C charge deposits mass of Cu=63.5 gram ∴ 1260 C charge will deposit mass of Cu= 63.5gram/96500x2 C X 1260 C =0.414 gram (1)	3
Q.72	The cell reaction is- Mg(s) + $Cu^{2+}_{(aq)}$----- > $Mg^{2+}_{(aq)}$ + Cu (s) here n = 2 (0.5)	2

	We know - $E^0_{cell} = E^0_{red.}(R) - E^0_{red.}(L)$ $E^0_{cell} = 0.34V - (-2.37V)$ $E^0_{cell} = 2.71V$ (0.5) Calculating equilibrium constant(K_c) $E^0 = 0.0591/n \times \log K_c$ ($E^0 = 2.303RT/nF$ at 298 K) $2.71 = 0.0591/2 \times \log K_c$ $\log K_c = 2.71 \times 2/0.0591$ $\log K_c = 91.7089$ (0.5) K_c = antilog 91.7089 (Take antilog $n.x = 0.x \times 10^n$) $K_c = 5.116 \times 10^{91}$ (0.5)	
Q.73	-Aman should use the Copper spoon. (0.5) -The reduction potential of Aluminium is lower than Fe metal hence Aluminium spoon (metal) cannot be used as it undergoes oxidation i.e.it would lose e^-. $Al_{(s)} \text{----> } Al^{3+}_{(aq)} + 3e^-$ (0.5) Thus, the solution will slowly turn into Aluminium sulphate and it will not serve the purpose. -Copper, on the other hand, has a reduction potential higher than Fe (both in ionic state) or we can say it is less reactive than Fe. $Cu^{2+} + 2e^- \text{----> } Cu_{(s)}$ (0.5) -Moreover, Cu here is in the solid state. Hence Cu spoon can be used to stir the solution as it will not bring any change in the $FeSO_4$ solution. (0.5)	2
Q.74	i) $Cu_{(s)}/Cu^{2+}_{(aq)}$ II $Ag^+_{(aq)}/Ag_{(s)}$ (1) ii) $\Delta G^0 = -n E^0 F$ n=2, $E^0 = 0.46V$ F= 96500C $\Delta G^0 = -2 \times 0.46V \times 96500C/mol$ $\Delta G^0 = -88780$ C.V/mol or -88780 J/mol (1C.V=1 Joule) or	2

	ΔG^0= - 88.78 kJ/mol (1)	
Q.75	A chemical reaction is feasible if E^0_{cell} is positive i.e. if potential of the cell is positive. We know- $E^0 = E^0_{red.(C)} - E^0_{red.(A)}$ (0.5) In the given equation ; $Ag_{(s)}$ ----> $Ag^+_{(aq.)} + e^-$ (oxidation) $Fe^{3+}_{(aq)} + 3e^-$ ---> $Fe_{(s)}$ (Reduction) (0.5) $E^0_{(cell)}$= 0.77V - 0.80V (values given) $E^0_{(cell)}$= - 0.03V (0.5) Since $E^0_{(cell)}$ is negative, reaction is not feasible. (0.5)	2
Q.76	For SHE: $H^+ + e$--- >$1/2H_2$ Applying Nernst eq. $E = E^0 - (0.0591V/n) (\log 1/[H^+])$ (1) $E = 0 - (0.0591V/1) (\log 1/1x10^{-10})$ E = 0.0591V x 10 log 10 E = 0.591V (log10=1) (1)	2
Q.77	- Charge required for the Reduction of 1 mole of $Cr_2O_7^{2-}$ ion is $Cr_2O_7^{2-} + 6e^- + H^+$---> $2Cr^{3+}$ 6 Faraday (0.5) - Therefore, charge required for 0.5 mole ion is $Cr_2O_7^{2-} + 3e^- + H^+$---> Cr^{3+} 3 Faraday (0.5) i.e. 96500C x 3 = 289500C (1 F= 96500C) (1)	2
Q.78	We know 1 Faraday electric charge deposits 1g equivalent of any substance on the electrodes kept in a series (According to Faraday's second law) Equivalent mass of Ag = 108/1 = 108g Equivalent mass of Cu = 63.5g/2 = 31.75g	2

	Equivalent mass of A l= 27/3 = 9.0g (1) Hence simple mass ratio deposited at respective electrodes are Ag : Cu : Al 108g : 31.75g : 9.0g 12 : 4 : 1 (Cu = 3.527 ≈ 4) (1)	
Q.79	i) In experimental Set up I, the blue colour of $CuSO_4$ solution will fade away. It is because $CuSO_4$ solution will turn into H_2SO_4 solution. Oxidation of water leaves behind H^+ and reduction of Cu^{2+} ion leaves SO_4^{2-} ion in the solution. $2H^+ + SO_4^{2-} \text{-----} > H_2SO_4$ ii) $Cu_{(s)} \text{-----} > Cu^{2+}_{(aq)} + 2e^-$ iii)Oxygen (O_2) ($2OH \text{-----} > O_2 + 2H^+ + 4e^-$) iv) Set up II depict the refining of Cu metal. In this setup, an impure copper rod is made anode, where oxidation takes place, At anode- $Cu_{(s)} \text{-----} > Cu^{2+}_{(aq)} + 2e^-$ and a pure thin wire of copper is made cathode. At cathode- $Cu^{2+}_{(aq)} + 2e \text{----} > Cu(s)$	3